Bohr Model of the Hydrogen Atom

Unit: Quantum and Particle Physics

NGSS Standards/MA Curriculum Frameworks (2016): N/A

AP® Physics 2 Learning Objectives/Essential Knowledge (2024): 15.2.A, 15.2.A.3, 15.2.A.3.i, 15.2.A.3.ii

Mastery Objective(s): (Students will be able to…)

- Explain how Bohr's model unified recent developments in the fields of spectroscopy, atomic theory and early quantum theory.
- Calculate the energy associated with a quantum number using Bohr's equation.

Success Criteria:

- Descriptions & explanations account for observed behavior.
- Variables are correctly identified and substituted correctly into the correct equation.
- Algebra is correct with correct rounding and reasonable units.

Language Objectives:

- Explain why the Bohr Model was such a big deal.

Tier 2 Vocabulary: model, quantum

Notes:

Significant Developments Prior to 1913

Atomic Structure

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Discovery of the Electron (1897): English physicist J.J. Thomson determined that cathode rays were actually particles emitted from atoms that the cathode was made of. Thomson called these particles "corpuscles", but eventually they came to be called "electrons" because they are "particles of electricity".

Saturnian Model of the Atom (1903): Japanese physicist Hantaro Nagaoka first proposed a model of the atom in which a small nucleus was surrounded by a ring of electrons. (Nagaoka actually used the term "electrons".) Other parts of his theory were disproved, and Nagaoka abandoned the theory in its entirety in 1908.

Discovery of the Atomic Nucleus (1906–1911): English physicists Hans Geiger and Ernest Marsden working under the direction Ernest Rutherford at the University of Manchester performed a series of scattering experiments that involved passing a beam of charged alpha particles through a thin sheet of gold foil. Most of the particles passed through unimpeded, but some were deflected, and a few were deflected sharply. These experiments confirmed the hypothesis that atoms contain a dense, positively charged nucleus that comprises most of the atom's mass, which is surrounded by negatively-charged electrons.

Early Quantum Theory

quantum: an elementary unit of energy.

In 1900, German physicist Max Planck published the Planck postulate, stating that electromagnetic energy could be emitted only in quantized form, *i.e.,* only certain "allowed" energy states are possible.

Plack determined the constant that bears his name as the relationship between the frequency of one quantum unit of electromagnetic wave and its energy. This relationship is the equation:

$$E = hf$$

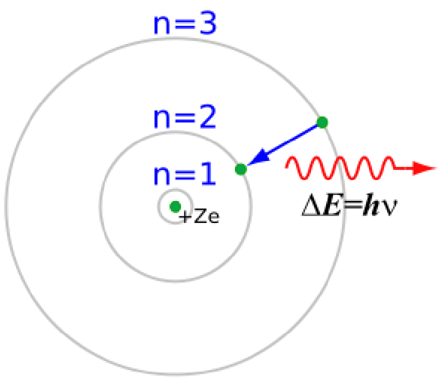
where:

E = energy (J)
h = Planck's constant = $6.626\times 10^{-34}\ \mathrm{J\cdot s}$
f = frequency* ($\mathrm{Hz} \equiv \mathrm{s}^{-1}$)

* Most physics texts use the Greek letter *ν* (nu) as the variable for frequency. However, high school texts and the College Board use *f*, presumably to avoid confusion with the letter "v".

Bohr's Model of the Atom (1913)

In 1913, Danish physicist Niels Bohr combined atomic theory, spectroscopy, and quantum theory into a single unified theory. Bohr hypothesized that electrons moved around the nucleus as in Rutherford's model, but that these electrons had only certain allowed quantum values of energy, which could be described by a quantum number (n). The value of that quantum number was the same n as in Rydberg's equation, and that using quantum numbers in Rydberg's equation could predict the wavelengths of light emitted when the electrons gained or lost energy by moving from one quantum level to another.



Bohr's model gained wide acceptance, because it related several prominent theories of the time. He received a Nobel Prize in physics in 1922 for his work.

The Bohr model is often given short shrift in high school chemistry classes, because it has been superseded by modern quantum theory. However, because Bohr's model unified several fields within physics, it is perhaps one of the most pivotal theories of the early 20th century.

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Bohr defined the energy associated with a given quantum number (n) in terms of Rydberg's constant:

$$E_n = -\frac{R_H}{n^2}$$

Although the model worked well for hydrogen, the equations could not be solved exactly for atoms with more than one electron, because of the additional effects that electrons exert on each other (*e.g.,* the Coulomb force, $F_e = \frac{1}{4\pi\varepsilon_o}\frac{q_1 q_2}{r^2}$).